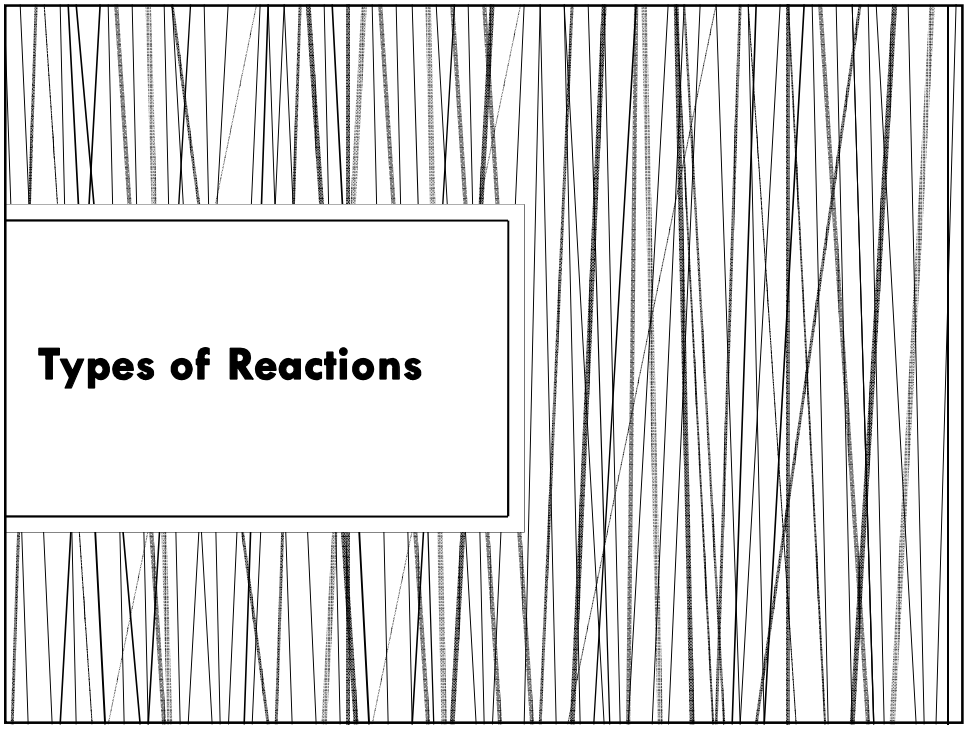

Chemical Equations

- Molecular equation—a chemical equation in which the formula for each substance is written without regard for whether it is an electrolyte or nonelectrolyte
  - $Pb(NO_3)_2\ (aq) + 2\ KI\ (aq) \rightarrow PbI_2\ (s) + 2\ KNO_3\ (aq)$
- Complete ionic equation—a chemical equation in which dissolved strong electrolytes (such as dissolved ionic compounds) are written as separate ions
  - $Pb^{2+}\ (aq) + 2\ NO_3^-\ (aq) + 2\ K^+\ (aq) + 2\ I^-\ (aq) \rightarrow PbI_2\ (s) + 2\ K^+\ (aq) + 2\ NO_3^-\ (aq)$
- Net ionic equation—a chemical equation for a solution reaction in which soluble strong electrolytes are written as ions and spectator ions are omitted
  - $Pb^{2+}\ (aq) + 2\ I^-\ (aq) \rightarrow PbI_2\ (s)$

Synthesis

- Also known as combination
- Two elements/compounds becoming one
- $A + X \rightarrow AX$
- Example:
  - $Fe\ (s) + Cl_2\ (g) \rightarrow FeCl_2$
- Example (metal oxide + water):
  - $Li_2O + H_2O \rightarrow 2\ LiOH$

Synthesis

- Example (nonmetal oxide + water):
  - $SO_3 + H_2O \rightarrow H_2SO_4$
- Example (metal oxide + nonmetal oxide):
  - $6\ CaO + P_4H_{10} \rightarrow 2\ Ca_3(PO_4)_2$

Decomposition

- Chemical breakdown through electrolysis or heat
- $AB \rightarrow A + B$
- Example (Electrolysis):
  - $2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$

Decomposition

- Example (Heating):
  - With <u>hydroxides</u>
    - $Ba(OH)_2 \rightarrow BaO + H_2O$
  - With <u>carbonates</u>
    - $BaCO_3 \rightarrow BaO + CO_3$
  - With <u>sulfates</u>
    - $FeSO_4 \rightarrow FeO + SO_3$

Decomposition

- Example (Heating):
  - With nitrates
    - $Ba(NO_3)_2 \rightarrow BaO + O_2 + NO_2$
  - With chlorates
    - $2KClO_3 \rightarrow 2KCl + 3O_2$
  - With oxides
    - $2HgO \rightarrow 2Hg + O_2$

Combustion

- Reaction of compound with oxygen gas
- Indicated by "burning" phrase in question
- $(C_xH_y) + O_2 \rightarrow CO_2 + H_2O$ **or**
  Metal (M) + $O_2 \rightarrow MO$
- Example (hydrocarbon):
  - $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
- Example (metal):
  - $2Mg + O_2 \rightarrow 2MgO$

Oxidation-reduction

- Also known as redox
- $A + BX \rightarrow AX + Y$ **or** $AX + Y \rightarrow AY + X$
- Use the activity series to predict products
- Example:
  - $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$
- Oxidation—increase in oxidation state
- Reduction—decrease in oxidation state
- Oxidizing agent—electron acceptor
- Reducing agent—electron donor

Oxidation States

- Imaginary charges the atoms would have if the shared electrons were equally between identical atoms bonded to each other
- Rules:
  - All atoms in an element is zero (Na, O_2, Hg)
  - A monatomic ion Is the same as its charge (Na^+, Cl^-)
  - Fluorine is -1 in its compounds
  - Oxygen is usually -2 in its compounds
  - Hydrogen is +1 in its covalent compounds

Common Reducing Agents

- Alkali metals → cation form
  - $4\ Li + O_2 \rightarrow 2\ Li_2O$
- Most metal/metallic ions → highest oxidation state of metal
  - $Cu + Cl_2 \rightarrow 2\ CuCl_2$
- Hydrogen → H^+
  - $H_2 + Cl_2 \rightarrow 2\ HCl$
- Carbon monoxide (CO) → Carbon dioxide (CO_2)
  - $2\ CO + O_2 \rightarrow 2\ CO_2$
- $H_2S + S^{2-} \rightarrow S$ (with all oxidizing agents except conc. HNO_3)
  - $3\ H_2S + K_2Cr_2O_7 + 4\ H_2SO_4 \rightarrow 3\ S + Cr_2(SO_4)_3 + K_2SO_4 + 7\ H_2O$
- $H_2S + S^{2-} \rightarrow SO_4^{2-}$ (with conc. HNO_3)
  - $H_2S + 8\ HNO_3$ (conc.) $\rightarrow H_2SO_4 + 2\ NO_2 + 4\ H_2O$

Common Reducing Agents

- SO_2 or $SO_3^{2-} \rightarrow SO_4^{2-}$
  - $2\ KMnO_4 + SO_2 + 2\ H_2O \rightarrow K_2SO_4 + 2\ MnSO_4 + H_2SO_4$
- Halides (Cl^-, Br^-, I^-) → Free halogens (Cl_2, Br_2, I_2)
  - $MnO_4 + 4\ HCl \rightarrow Cl_2 + MnCl_2 + 2\ H_2O$
- $HNO_2 \rightarrow HNO_3$
  - $5\ HNO_2 + 2\ KMnO_4 + 3\ H_2SO_4 \rightarrow 5\ HNO_3 + 2\ MnSO_4 + K_2SO_4 + 3\ H_2O$
- $H_2O_2 \rightarrow O_2$
  - $5\ H_2O_2 + 2\ KMnO_4 + 3\ H_2SO_4 \rightarrow K_2SO_4 + 2\ MnSO_4 + 5\ O_2 + 8\ H_2O$

Common Oxidizing Agents

- $O_2 \rightarrow O^{2-}$ (sometimes H_2O)
  - $2\ Ca + O_2 \rightarrow 2\ CaO$
- Halogens (Cl_2, Br_2, I_2) $\rightarrow$ Halides (Cl^-, Br^-, I^-)
  - $Zn + Br_2 \rightarrow ZnBr$
- Dilute $HNO_3 \rightarrow NO_2$
  - $3\ Cu + 8\ HNO_3\ (dilute) \rightarrow 3\ Cu(NO_3)_2 + 2\ NO + 4\ H_2O$
- Concentrated $HNO_3 \rightarrow NO_2$
  - $Cu + 4\ HNO_3\ (conc.) \rightarrow 3\ Cu(NO_3)_2 + 2\ NO_2 + 2\ H_2O$
- MnO_4 or $MnO_2 \rightarrow Mn^{2+}$ (in acidic solution)
  - $MnO_2 + 4\ HCl \rightarrow MnCl_2 + Cl_2 + 2\ H_2O$
- $MnO_4 \rightarrow MnO_2$ (in basic or neutral solution)
  - $KMnO_4 + 3\ NO_2 + 2\ KOH \rightarrow 3\ KNO_3 + 2\ MnO_2 + H_2O$

Common Oxidizing Agents

- $Cr_2O_7^{2-}$ or $CrO_4^{2-} \rightarrow Cr^{3+}$
  - $K_2Cr_2O_7 + 14\ HCl \rightarrow 2\ CrCl_3 + 2\ KCl + 3\ Cl_2 + 7\ H_2O$
- $HNO_2 \rightarrow NO$
  - $2\ HNO_2 + 2\ HI \rightarrow I_2 + 2\ NO + 2\ H_2O$
- $H_2O_2 \rightarrow H_2O$
  - $2\ HI + H_2O_2 \rightarrow 2\ H_2O + I_2$

Double Displacement—Precipitation

- A precipitate forms from two soluble reactants
- Use solubility rules to predict products
- AX + BY → BX + AY
- Example:
  - $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

Double Displacement—Acid-Base

- Acid—substance that forms hydrogen ions
- Base—substance that accepts hydrogen ions
- Only neutralization reactions
- AX + BY → AY + BX
- Example (acid + base):
  - $HCl + KOH \rightarrow H_2O + KCl$
- Example (Metallic oxide + acid):
  - $BaO + 2\ HCl \rightarrow H_2O + BaCl_2$
- Example (Nonmetallic oxide + base):
  - $SO_3 + 2\ KOH \rightarrow H_2O + K_2SO_4$

Strong Acids & Bases

- Strong Acids
  - HCl
  - HBr
  - HI
  - $HClO_3$
  - $HClO_4$
  - HNO_3
  - H_2SO_4
- Strong Bases
  - LiOH
  - NaOH
  - KOH
  - RbOH
  - CsOH
  - $Ca(OH)_2$
  - $Sr(OH)_2$
  - $Ba(OH)_2$